Electron Configuration

Quantum Numbers:

Name	**Symbol**	**Possible values**	**Defines**
Principal Quantum Number	n	1 to infinity	Energy level and size of orbital
Azimuthal Quantum Number	l	0 to (n-1)	Shape of orbital
Magnetic Quantum Number	m_l	$-l$ to l	Orientation
Spin Quantum Number	m_s	+1/2, -1/2	Electron spin

Electron Configuration

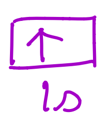
- The __arrangement__ of electrons in the orbitals.
- For example, consider the hydrogen electron:
    - It is most likely to be found in the __1s__ orbital.
    - Its ground-state electron configuration would be: __$1s^1$__.

      # of electrons

      energy level (n)

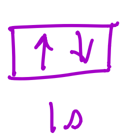
      orbital shape (ℓ)

- We can accompany an electron configuration for an atom with an __orbital box__.
    - Each orbital in any given principal energy level is represented as a __box__.
    - Each __box__ can contain either 0, 1, or 2 electrons, which we represent by an arrow.
        - ➢ Arrow pointing __up (↑)__ to represent an electron with a spin of +1/2.
        - ➢ Arrow pointing __down (↓)__ to represent an electron with a spin of -1/2.
    - For example, the orbital diagram that would accompany the ground-state configuration of the hydrogen atom in the previous example would be:



$1s^1$

- The orbital diagram of helium that would accompany its ground-state electron configuration of __$1s^2$__ would be:



$1s^2$

- Condensed electron configuration and orbital diagrams
  - Used to reduce the length of the electron configuration.
  - Use the atomic symbol of the ___nobel___ gas from the previous period in brackets, which will represent that gas' electron configuration.
  - For example, the electron configuration of a ~~sodium~~ magnesium atom is $1s^2 2s^2 2p^6 3s^2$.
    - ➢ The nobel gas from the previous period is neon (Ne) with electron configuration $1s^2 2s^2 2p^6$.
    - ➢ The condensed electron configuration for a ~~sodium~~ magnesium atom is: ___$[Ne]3s^2$___.
  - When drawing the orbital diagram of a condensed electron configuration, it is enough to only represent the orbital diagrams for the ___Valence___ electrons.

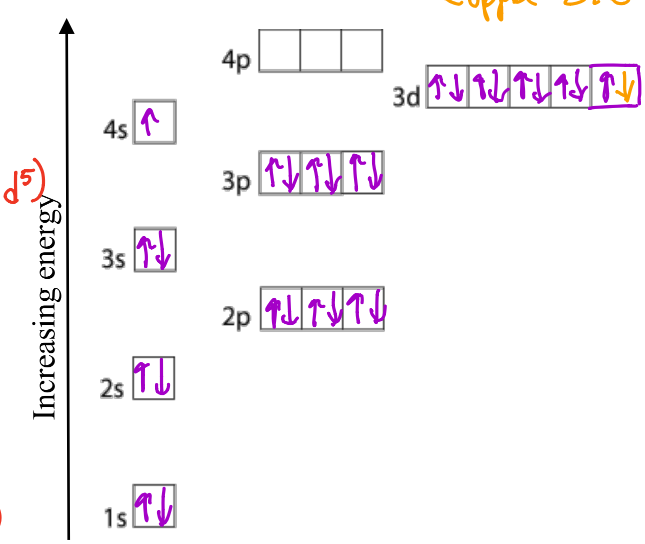
$[Ne]3s^2$: $\boxed{\uparrow\downarrow}$ 3s

- When writing electron configurations and orbital diagrams, it is important that we follow the following principles and rule:
  - <u>Aufbau principle:</u>
    - ➢ Electrons will always fill the ___lowest___ available energy level.
    - ➢ The diagram below can be useful in identifying the lowest available energy level (however, ***<u>there are exceptions</u>***!)
    - ➢ Chromium:

$[Ar]4s^2 3d^4$

Reality $[Ar]4s^1 3d^5$

– Forming a half-full sublevel (i.e, d^5)



- ➢ Copper

$[Ar]4s^2 3d^9$

Reality $[Ar]4s^1 3d^{10}$

– Forming a full level ($3s^2\ 3p^6\ 3d^{10}$)

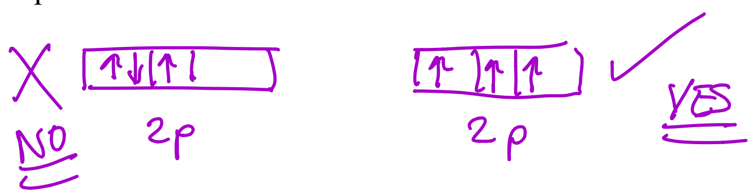
- ➢ Takeaways from exceptions:

① Better to have a full level (n) *****

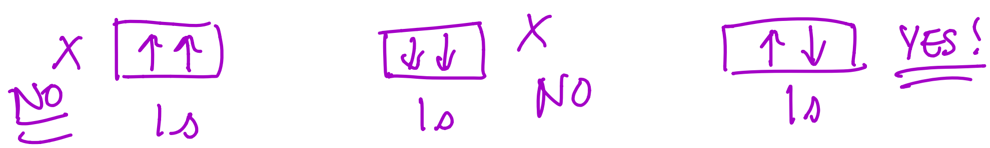
② Better to have a full sublevel (ℓ) ***

③ Better to have a half full sublevel (ℓ) **

- Hund's rule
  - Each orbital must be occupied by a single electron until it is no longer possible to do so.



- Pauli exclusion principle:
  - Each orbital can only contain a maximum of two electrons, and they must be of ___opposite___ spins.
  - Example:



- Using the periodic table to predict electron configuration
  - Groups 1 and 2: ___s-block elements___.
  - Groups 13 to 18: ___p-block elements___.
  - Group 12 and the transition elements: ___d-block elements___.
  - Inner transition elements (lanthanides) and 89-103 (actinides): ___f-block elements___.